

# Chemical Reactions

## What You'll Learn

- ▶ You will write chemical equations to describe chemical reactions.
- ▶ You will classify and identify chemical reactions.
- ▶ You will write ionic equations for reactions that occur in aqueous solutions.

## Why It's Important

Chemical reactions affect you every second of every day. For example, life-sustaining chemical reactions occur continuously in your body. Other chemical reactions occur in less likely situations, such as in a thunderstorm.



CLICK HERE

Visit the Chemistry Web site at [science.glencoe.com](http://science.glencoe.com) to find links about chemical reactions.

The electricity of a lightning bolt provides the energy that sparks chemical reactions among substances in the atmosphere.



## DISCOVERY LAB



### Materials

distilled water	universal indicator
25-mL graduated cylinder	stirring rod
100-mL beaker	thermometer
pipettes (2)	effervescent tablet
0.1M ammonia	

### Observing a Change

An indicator is a chemical that shows when change occurs during a chemical reaction.

### Safety Precautions



Always wear goggles and an apron in the laboratory.

### Procedure

1. Measure 10.0 mL distilled water in a graduated cylinder and pour it into the beaker. Add one drop of 0.1M ammonia to the water.
2. Stir 15 drops of indicator into the solution with the stirring rod. Observe the solution's color. Measure its temperature with the thermometer.
3. Drop the effervescent tablet into the solution. Observe what happens. Record your observations, including any temperature change.

### Analysis

Did a color change and a temperature change occur? Was a gas produced? Did a physical change or a chemical change occur? Explain.

## Section

## 10.1

# Reactions and Equations

### Objectives

- **Recognize** evidence of chemical change.
- **Represent** chemical reactions with equations.

### Vocabulary

chemical reaction  
reactant  
product  
chemical equation  
coefficient

Do you know that the foods you eat, the fibers in your clothes, and the plastic in your CDs have something in common? Foods, fibers, and plastic are produced when the atoms in substances are rearranged to form different substances. Atoms are rearranged during the flash of lightning shown in the photo on the opposite page. They were also rearranged when you dropped the effervescent tablet into the beaker of water and indicator in the **DISCOVERY LAB**.

### Evidence of Chemical Reactions

The process by which the atoms of one or more substances are rearranged to form different substances is called a **chemical reaction**. A chemical reaction is another name for a chemical change, which you read about in Chapter 3. Chemical reactions affect every part of your life. They break down your food, producing the energy you need to live. They produce natural fibers such as cotton and wool in the bodies of plants and animals. In factories, they produce synthetic fibers such as nylon and polyesters. Chemical reactions in the engines of cars and buses provide the energy to power the vehicles.

How can you tell when a chemical reaction has taken place? Although some chemical reactions are hard to detect, many reactions provide evidence that they have occurred. A temperature change can indicate a chemical reaction. Many reactions, such as those that occur during a forest fire, release energy in the form of heat and light. Other reactions absorb heat.

In addition to a temperature change, other types of evidence may indicate that a chemical reaction has occurred. One indication of a chemical reaction is

**Figure 10-1**

Each of these photos illustrates evidence of a chemical reaction.

- a** Reactions that happened when the marshmallow was burned are obvious by the color change.
- b** Chemical reactions occur in the oven when a cake mix is baked, namely, the formation of gas bubbles that cause the cake to rise.
- c** The tarnish that appears on silver and other metals is actually a solid that forms as a result of chemical reactions that take place when the metal is exposed to traces of sulfur compounds in the air.
- d** Numerous chemical reactions happen in an explosion. The appearance of smoke, the release of energy in the form of heat, and the permanent color change of materials involved are all evidence of chemical reactions.



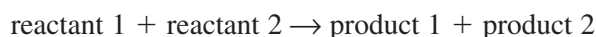
a color change. For example, you may have noticed that the color of some nails that are left outside changes from silver to orange-brown in a short time. The color change is evidence that a chemical reaction occurred between the iron in the nail and the oxygen in air. Odor, gas bubbles, and/or the appearance of a solid are other indications of chemical change. Each of the photographs in **Figure 10-1** shows evidence of a chemical reaction. Do you recognize the evidence in each?

**Table 10-1**

Symbols Used in Equations	
Symbol	Meaning
+	Separates two or more reactants or products
→	Separates reactants from products
(s)	Identifies solid state
(l)	Identifies liquid state
(g)	Identifies gaseous state
(aq)	Identifies water solution

## Representing Chemical Reactions

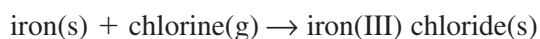
Chemists use statements called equations to represent chemical reactions. Their equations show a reaction's **reactants**, which are the starting substances, and **products**, which are the substances formed during the reaction. Chemical equations do not express numerical equalities as do mathematical equations because during chemical reactions the reactants are used up as the products form. Instead, the equations used by chemists show the direction in which the reaction progresses. Therefore, an arrow rather than an equal sign is used to separate the reactants from the products. You read the arrow as “react to produce” or “yield”. The reactants are written to the arrow's left, and the products are written to its right. When there are two or more reactants, or two or more products, a plus sign separates each reactant or each product. These elements of equation notation are shown below.





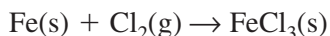
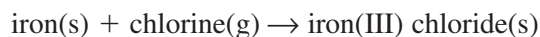
In equations, symbols are used to show the physical states of the reactants and products. Reactants and products can exist as solids, liquids, and gases. When they are dissolved in water, they are said to be aqueous. It is important to show the physical states of a reaction's reactants and products in an equation because the physical states provide clues about how the reaction occurs. Some basic symbols used in equations are shown in **Table 10-1**.

**Word equations** You can use statements called word equations to indicate the reactants and products of chemical reactions. The word equation below describes the reaction between iron and chlorine, which is shown in **Figure 10-2**. Iron is a solid and chlorine is a gas. The brown cloud in the photograph is composed of the reaction's product, which is solid particles of iron(III) chloride.



This word equation is read, "Iron and chlorine react to produce iron(III) chloride."

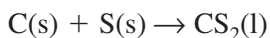
**Skeleton equations** Although word equations help to describe chemical reactions, they are cumbersome and lack important information. A skeleton equation uses chemical formulas rather than words to identify the reactants and the products. For example, the skeleton equation for the reaction between iron and chlorine uses the formulas for iron, chlorine, and iron(III) chloride in place of the words.



How would you write the skeleton equation that describes the reaction between carbon and sulfur to form carbon disulfide? Carbon and sulfur are solids. First, write the chemical formulas for the reactants to the left of an arrow. Then, separate the reactants with a plus sign and indicate their physical states.



Finally, write the chemical formula for the product, liquid carbon disulfide, to the right of the arrow and indicate its physical state. The result is the skeleton equation for the reaction.



This skeleton equation tells us that the reaction of carbon in the solid state reacts with sulfur in the solid state to produce carbon disulfide, which is in the liquid state.



**Figure 10-2**

Science, like all other disciplines, has a specialized language that allows specific information to be communicated in a uniform manner. This reaction between iron and chlorine can be described by a word equation, skeleton equation, or balanced chemical equation.

## PRACTICE PROBLEMS

Write skeleton equations for the following word equations.

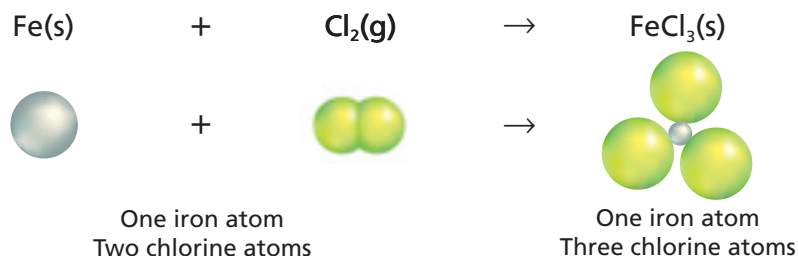
1. hydrogen(g) + bromine(g)  $\rightarrow$  hydrogen bromide(g)
2. carbon monoxide(g) + oxygen(g)  $\rightarrow$  carbon dioxide(g)
3. potassium chlorate(s)  $\rightarrow$  potassium chloride(s) + oxygen(g)



For more practice with writing skeleton equations, go to **Supplemental Practice Problems** in Appendix A.

**Figure 10-3**

The information conveyed by skeleton equations is limited. In this case, the skeleton equation (top) is correct, but it does not show the exact number of atoms that actually interact. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



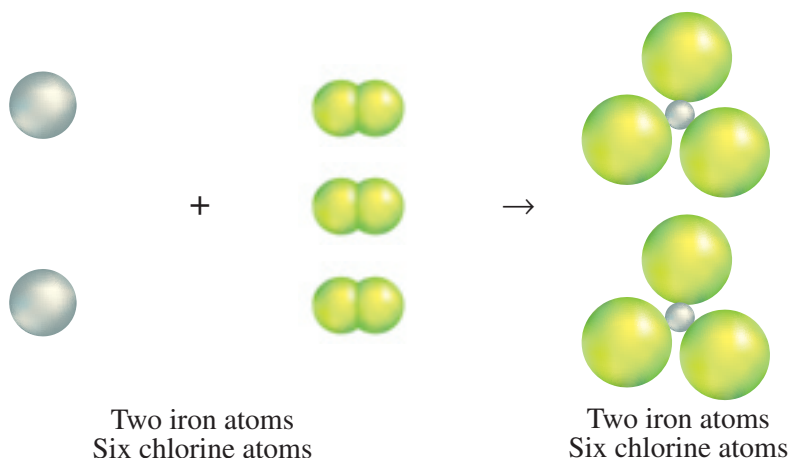
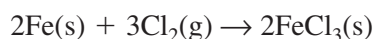
**Chemical equations** Writing a skeleton equation is an important step toward using an equation to completely describe a chemical reaction. But, like word equations, skeleton equations also lack important information about reactions. Recall from Chapter 3 that the law of conservation of mass states that in a chemical change, matter is neither created nor destroyed. Chemical equations must show that matter is conserved during a reaction, and skeleton equations lack that information.

Look at **Figure 10-3**. The skeleton equation for the reaction between iron and chlorine shows that one iron atom and two chlorine atoms react to produce a substance containing one iron atom and three chlorine atoms. Was a chlorine atom created in the reaction? Atoms are not created in chemical reactions, and to accurately show what happened, more information is needed.

To accurately represent a chemical reaction by an equation, the equation must show how the law of conservation of mass is obeyed. In other words, the equation must show that the number of atoms of each reactant and each product is equal on both sides of the arrow. Such an equation is called a balanced **chemical equation**. A chemical equation is a statement that uses chemical formulas to show the identities and relative amounts of the substances involved in a chemical reaction. It is chemical equations that chemists use most often to represent chemical reactions.

## Balancing Chemical Equations

The balanced equation for the reaction between iron and chlorine, shown below, reflects the law of conservation of mass.



To balance an equation, you must find the correct coefficients for the chemical formulas in the skeleton equation. A **coefficient** in a chemical equation is the number written in front of a reactant or product. Coefficients are usually whole numbers, and are usually not written if the value is 1. A coefficient

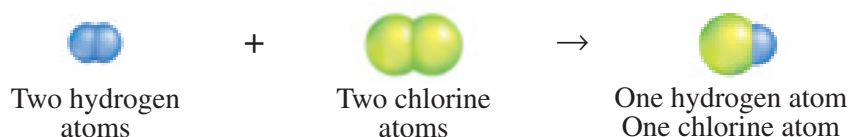
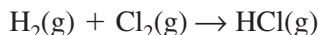


Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

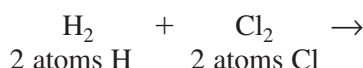
tells you the smallest number of particles of the substance involved in the reaction. That is, the coefficients in a balanced equation describe the lowest whole-number ratio of the amounts of all of the reactants and products.

**Steps for balancing equations** Most chemical equations can be balanced by following the steps given below. For example, you can use these steps to write the chemical equation for the reaction between hydrogen and chlorine that produces hydrogen chloride.

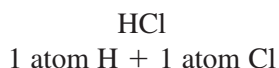
**Step 1** Write the skeleton equation for the reaction. Make sure that the chemical formulas correctly represent the substances. An arrow separates the reactants from the products, and a plus sign separates multiple reactants and products. Show the physical states of all reactants and products.



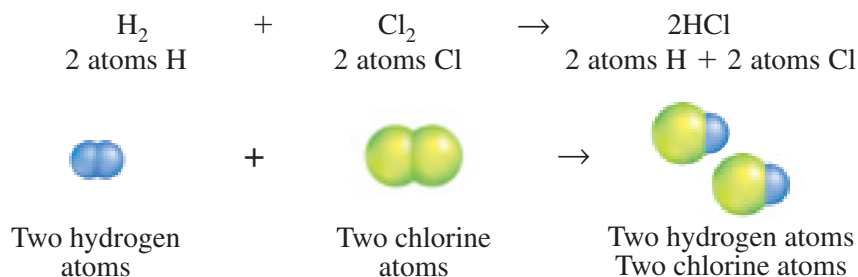
**Step 2** Count the atoms of the elements in the reactants. If a reaction involves identical polyatomic ions in the reactants and products, count the ions as if they are elements. This reaction does not involve any polyatomic ions. Two atoms of hydrogen and two atoms of chlorine are reacting.



**Step 3** Count the atoms of the elements in the products. One atom of hydrogen and one atom of chlorine are produced.



**Step 4** Change the coefficients to make the number of atoms of each element equal on both sides of the equation. Never change a subscript in a chemical formula to balance an equation because doing so changes the identity of the substance.



**Step 5** Write the coefficients in their lowest possible ratio. The coefficients should be the smallest possible whole numbers. The ratio 1 hydrogen to 1 chlorine to 2 hydrogen chloride (1:1:2) is the lowest possible ratio because the coefficients cannot be reduced and still remain whole numbers.

**Step 6** Check your work. Make sure that the chemical formulas are written correctly. Then, check that the number of atoms of each element is equal on both sides of the equation.

## Earth Science

### CONNECTION

**W**eathering is the general term used to describe the ways in which rock is broken down at or near Earth's surface. Soils are the result of weathering and the activities of plants and animals.

Physical weathering, also called mechanical weathering, involves expansion and contraction with changes in temperature, pressure, and the growth of plants and organisms in the rock. Water in rock fissures and crevices cause rock to fracture when water expands during freezing. Freeze-thaw physical weathering is more likely to occur in sub-Arctic climates.

Chemical weathering involves the break down of rock by chemical reactions. The mineral composition of the rock is changed, reorganized, or redistributed. For example, minerals that contain iron may react with oxygen in the air. Water in which carbon dioxide is dissolved will dissolve limestone. Chemical weathering is more likely to take place in humid tropical climates.



## EXAMPLE PROBLEM 10-1

### Writing a Balanced Chemical Equation

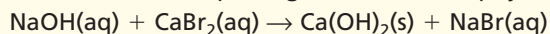
Write the balanced chemical equation for the reaction in which sodium hydroxide and calcium bromide react to produce solid calcium hydroxide and sodium bromide. The reaction occurs in water.

#### 1. Analyze the Problem

You are given the reactants and products in a chemical reaction. Start with a skeleton equation and use the steps given in the text for balancing chemical equations.

#### 2. Solve for the Unknown

*Step 1* Write the skeleton equation. Be sure to put the reactants on the left side of an arrow and the products on the right. Separate the substances with plus signs and indicate physical states.



*Step 2* Count the atoms of each element in the reactants.

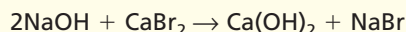
1 Na, 1 O, 1 H, 1 Ca, 2 Br

*Step 3* Count the atoms of each element in the products.

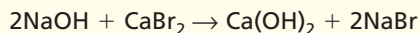
1 Na, 2 O, 2 H, 1 Ca, 1 Br

*Step 4* Adjust the coefficients.

Insert the coefficient 2 in front of NaOH to balance the hydroxide ions.



Insert the coefficient 2 in front of NaBr to balance the Na and Br atoms.



*Step 5* Write the coefficients in their lowest possible ratio.

The ratio of the coefficients is 2:1:1:2.

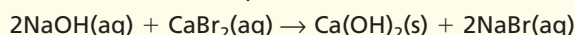
*Step 6* Check to make sure that the number of atoms of each element is equal on both sides of the equation.

Reactants: 2 Na, 2 OH, 1 Ca, 2 Br

Products: 2 Na, 2 OH, 1 Ca, 2 Br

#### 3. Evaluate the Answer

The chemical formulas for all substances are written correctly. The number of atoms of each element is equal on both sides of the equation. The coefficients are written in the lowest possible ratio. The balanced chemical equation for the reaction is



The brick and mortar used in many construction applications involves chemical reactions with calcium and other substances. Although these materials have been used for centuries, chemistry continues to improve their durability and performance.

## PRACTICE PROBLEMS

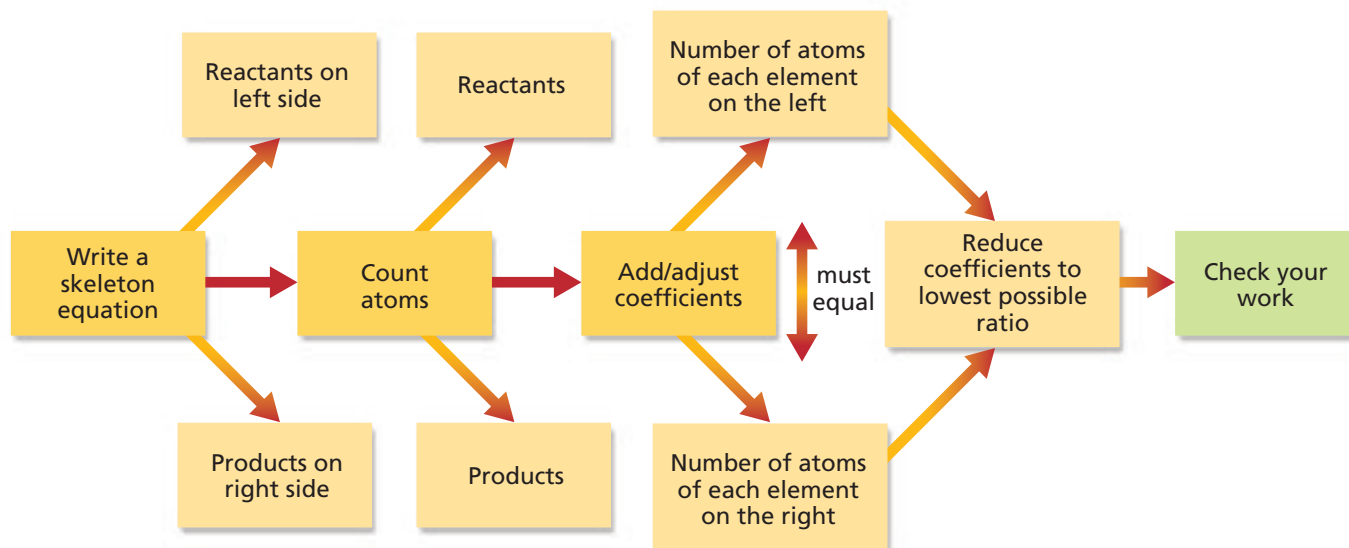
Write chemical equations for each of the following reactions.

4. In water, iron(III) chloride reacts with sodium hydroxide, producing solid iron(III) hydroxide and sodium chloride.
5. Liquid carbon disulfide reacts with oxygen gas, producing carbon dioxide gas and sulfur dioxide gas.
6. Solid zinc and aqueous hydrogen sulfate react to produce hydrogen gas and aqueous zinc sulfate.



For more practice writing chemical equations, go to **Supplemental Practice Problems** in Appendix A.

### Balancing Chemical Equations



**Figure 10-4**

It is imperative to your study of chemistry to be able to balance chemical equations. Use this flowchart to help you master the skill.

Probably the most fundamental concept of chemistry is the law of conservation of mass that you first encountered in Chapter 3. All chemical reactions obey the law that matter is neither created nor destroyed. Therefore, it is also fundamental that the equations that represent chemical reactions include sufficient information to show that the reaction obeys the law of conservation of mass. You have learned how to show this relationship with balanced chemical equations. The flowchart shown in **Figure 10-4** summarizes the steps for balancing equations. You will undoubtedly find that some chemical equations can be balanced easily, whereas others are more difficult to balance. All chemical equations, however, can be balanced by the process you learned in this section.

## Section 10.1 Assessment

- List three types of evidence that indicate a chemical reaction has occurred.
- Compare and contrast a skeleton equation and a chemical equation.
- Why is it important that a chemical equation be balanced?
- When balancing a chemical equation, can you adjust the number that is subscripted to a substance formula? Explain your answer.
- Why is it important to reduce coefficients in a balanced equation to the lowest possible whole-number ratio?
- Thinking Critically** Explain how an equation can be balanced even if the number of reactant particles differs from the number of product particles.
- Using Numbers** Is the following equation balanced? If not, correct the coefficients.  

$$2\text{K}_2\text{CrO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbCrO}_4(\text{s})$$



## Objectives

- **Classify** chemical reactions.
- **Identify** the characteristics of different classes of chemical reactions.

## Vocabulary

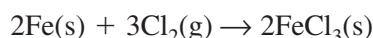
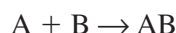
synthesis reaction  
combustion reaction  
decomposition reaction  
single-replacement reaction  
double-replacement reaction  
precipitate

How long do you think it would take you to find your favorite author's new novel in an unorganized book store? Because there are so many books in book stores, it could take you a very long time. Book stores, such as the store shown in **Figure 10-5**, supermarkets, and music stores are among the many places where things are classified and organized. Chemists classify chemical reactions in order to organize the many reactions that occur daily in living things, laboratories, and industry. Knowing the categories of chemical reactions can help you remember and understand them. It also can help you recognize patterns and predict the products of many chemical reactions.

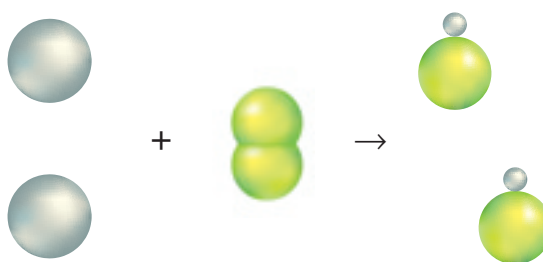
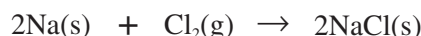
Chemists classify reactions in different ways. One way is to distinguish among five types of chemical reactions: synthesis, combustion, decomposition, single-replacement, and double-replacement reactions. Some reactions fit equally well into more than one of these classes.

## Synthesis Reactions

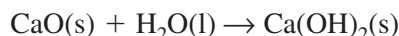
In the previous section, you read about the reaction that occurs between iron and chlorine gas to produce iron(III) chloride. In this reaction, two elements (A and B) combine to produce one new compound (AB).



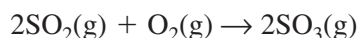
The reaction between iron and chlorine gas is an example of a **synthesis reaction**—a chemical reaction in which two or more substances react to produce a single product. When two elements react, the reaction is always a synthesis reaction. Another example of a synthesis reaction is shown below. In this reaction, sodium and chlorine react to produce sodium chloride.



Just as two elements can combine, two compounds can also combine to form one compound. For example, the reaction between calcium oxide and water to form calcium hydroxide is a synthesis reaction.



Another type of synthesis reaction may involve a reaction between a compound and an element, as happens when sulfur dioxide gas reacts with oxygen gas to form sulfur trioxide.



**Figure 10-5**

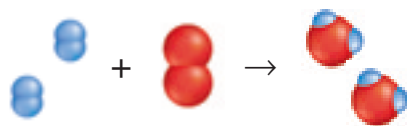
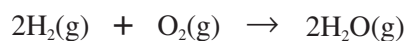
Without some level of organization, it would be difficult to shelve and maintain the huge number of books that have been published. Chemistry, too, hinges on strict rules of organization. Chemical reactions are classified as synthesis, combustion, decomposition, single-replacement, and double-replacement reactions.



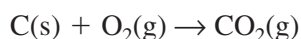
## Combustion Reactions

The synthesis reaction between sulfur dioxide and oxygen can be classified also as a combustion reaction. In a **combustion reaction**, oxygen combines with a substance and releases energy in the form of heat and light. Oxygen can combine in this way with many different substances, making combustion reactions common.

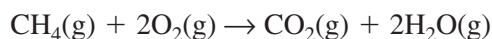
A combustion reaction, such as the one shown in **Figure 10-6**, occurs between hydrogen and oxygen when hydrogen is heated. Water is formed during the reaction and a large amount of energy is released.



Another important combustion reaction occurs when coal is burned to produce energy. Coal is called a fossil fuel because it contains the remains of plants that lived long ago. It is composed primarily of the element carbon. Coal-burning power plants generate electric power in many parts of the United States. The primary reaction that occurs in these plants is between carbon and oxygen.



Note that the combustion reactions just mentioned are also synthesis reactions. However, not all combustion reactions are synthesis reactions. For example, the reaction involving methane gas,  $\text{CH}_4$ , and oxygen illustrates a combustion reaction in which one substance replaces another in the formation of products.



Methane, which belongs to a group of substances called hydrocarbons, is the major component of natural gas. All hydrocarbons contain carbon and hydrogen and burn in oxygen to yield the same products as methane does—carbon dioxide and water. For example, most cars and trucks are powered by gasoline, which contains hydrocarbons. In engines, gasoline is combined with oxygen, producing carbon dioxide, water, and energy that powers the vehicles. You will learn more about hydrocarbons in Chapter 22.



**Figure 10-6**

The tragedy of the *Challenger* space mission was the result of a combustion reaction between oxygen and hydrogen.

### PRACTICE PROBLEMS

Write chemical equations for the following reactions. Classify each reaction into as many categories as possible.

14. The solids aluminum and sulfur react to produce aluminum sulfide.
15. Water and dinitrogen pentoxide gas react to produce aqueous hydrogen nitrate.
16. The gases nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.
17. Ethane gas ( $\text{C}_2\text{H}_6$ ) burns in air, producing carbon dioxide gas and water vapor.



For more practice with writing synthesis and combustion equations, go to **Supplemental Practice Problems** in Appendix A.

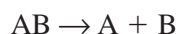
**Figure 10-7**

The decomposition of sodium azide, which produces a gas, is the key to inflatable air bags.



## Decomposition Reactions

Some chemical reactions are essentially the opposite of synthesis reactions. These reactions are classified as decomposition reactions. A **decomposition reaction** is one in which a single compound breaks down into two or more elements or new compounds. In generic terms, decomposition reactions look like the following.



Decomposition reactions often require an energy source, such as heat, light, or electricity, to occur. For example, ammonium nitrate breaks down into dinitrogen monoxide and water when the reactant is heated to high temperature.



You can see that this decomposition reaction involves one reactant breaking down into more than one product.

The outcome of another decomposition reaction is shown in **Figure 10-7**. Automobile safety air bags inflate rapidly as sodium azide pellets decompose. A device that can provide an electric signal to start the reaction is packaged inside air bags along with the sodium azide pellets. When the device is activated, sodium azide decomposes, producing nitrogen gas that quickly inflates the safety bag.



### PRACTICE PROBLEMS

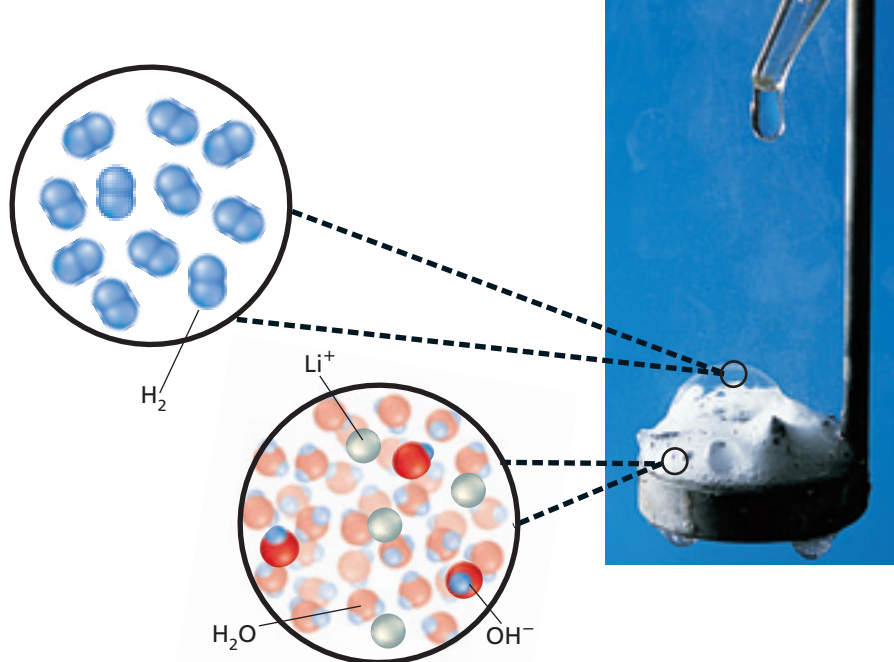
Write chemical equations for the following decomposition reactions.

18. Aluminum oxide(s) decomposes when electricity is passed through it.
19. Nickel(II) hydroxide(s) decomposes to produce nickel(II) oxide(s) and water.
20. Heating sodium hydrogen carbonate(s) produces sodium carbonate(aq), carbon dioxide(g), and water.

Practice!

For more practice with writing decomposition equations, go to

**Supplemental Practice Problems** in Appendix A.



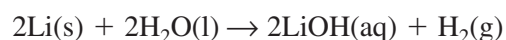
**Figure 10-8**

The reaction of lithium and water is a single-replacement reaction. Lithium replaces a hydrogen in water, and the products of the reaction are aqueous lithium hydroxide and hydrogen gas. Lithium hydroxide exists as lithium and hydroxide ions in solution.

## Replacement Reactions

In contrast to synthesis, combustion, and decomposition reactions, many chemical reactions involve the replacement of an element in a compound. There are two types of replacement reactions: single-replacement reactions and double-replacement reactions.

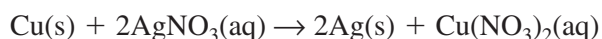
**Single-replacement reactions** Now that you've seen how atoms and molecules rearrange in synthesis and combustion reactions, look closely at the reaction between lithium and water that is shown in **Figure 10-8**. The expanded view of the reaction at the molecular level shows that a lithium atom replaces one of the hydrogen atoms in a water molecule. The following chemical equation describes this activity.



A reaction in which the atoms of one element replace the atoms of another element in a compound is called a **single-replacement reaction**.



The reaction between lithium and water is one type of single-replacement reaction in which a metal replaces a hydrogen in a water molecule. Another type of single-replacement reaction occurs when one metal replaces another metal in a compound dissolved in water. For example, **Figure 10-9** shows a single-replacement reaction occurring when a spiral of pure copper wire is placed in aqueous silver nitrate. The shiny crystals that are accumulating on the copper wire are the silver atoms that the copper atoms replaced.





A metal will not always replace another metal in a compound dissolved in water. This is because metals differ in their reactivities. A metal's reactivity is its ability to react with another substance. In **Figure 10-10** you see an activity series of some metals. This series orders metals by their reactivity with other metals. Single-replacement reactions like the one between copper and aqueous silver nitrate determine a metal's position on the list. The most active



**Figure 10-9**

The chemical equation for the single-replacement reaction involving copper and silver nitrate clearly describes the replacement of silver by copper, but the visual evidence of this chemical reaction is a solid precipitate.



Most active 	<b>METALS</b>
	Lithium
	Rubidium
	Potassium
	Calcium
	Sodium
	Magnesium
	Aluminum
	Manganese
	Zinc
	Iron
	Nickel
	Tin
	Lead
	Copper
	Silver
	Platinum
	Gold
Least active	
Most active 	<b>HALOGENS</b>
	Fluorine
	Chlorine
	Bromine
Least active	Iodine

**Figure 10-10**

An activity series, similar to the series shown here for various metals and halogens, is a useful tool for determining whether a chemical reaction will take place and for determining the result of a replacement reaction.

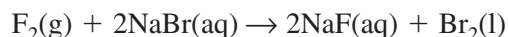
metals, which are those that do replace the metal in a compound, are at the top of the list. The least active metals are at the bottom.

You can use **Figure 10-10** to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, you saw in **Figure 10-9** that copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper(II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series and no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.



The **CHEMLAB** at the end of this chapter gives you an opportunity to explore the activities of metals in the laboratory.

A third type of single-replacement reaction involves the replacement of a nonmetal in a compound by another nonmetal. Halogens are frequently involved in these types of reactions. Like metals, halogens exhibit different activity levels in single-replacement reactions. The reactivities of halogens, determined by single-replacement reactions, are also shown in **Figure 10-10**. The most active halogen is fluorine, and the least active is iodine. A more reactive halogen replaces a less reactive halogen that is part of a compound dissolved in water. For example, fluorine replaces bromine in water containing dissolved sodium bromide. However, bromine does not replace fluorine in water containing dissolved sodium fluoride.



The **problem-solving LAB** below will help you to relate periodic trends of the halogens to their reactivities.

## problem-solving LAB

### Can you predict the reactivities of halogens?

**Analyzing and Concluding** The location of all the halogens in group 7A in the periodic table tells you that halogens have common characteristics. Indeed, halogens are all nonmetals and have seven electrons in their outermost orbitals. However, each halogen has its own characteristics, too, such as its ability to react with other substances.

#### Analysis

Examine the accompanying table. It includes data about the atomic radii, ionization energies, and electronegativities of the halogens.

#### Thinking Critically

1. Describe any periodic trends that you identify in the table data.

Properties of Halogens			
Halogen	Atomic radius (pm)	Ionization energy (kJ/mol)	Electronegativity
Fluorine	72	1681	3.98
Chlorine	100	1251	3.16
Bromine	114	1140	2.96
Iodine	133	1008	2.66
Astatine	140	—	2.2

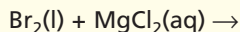
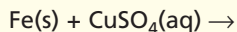
2. Relate any periodic trends that you identify among the halogens to the activity series of halogens shown in **Figure 10-10**.

3. Predict the location of the element astatine in the activity series of halogens. Explain your answer.

## EXAMPLE PROBLEM 10-2

### Single-Replacement Reactions

Predict the products that will result when these reactants combine and write a balanced chemical equation for each reaction.

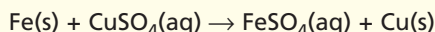


#### 1. Analyze the Problem

You are given three sets of reactants. Using **Figure 10-10**, you must first determine if each reaction takes place. Then, if a reaction is predicted, you can determine the product(s) of the reaction. With this information you can write a skeleton equation for the reaction. Finally, you can use the steps for balancing chemical equations to write the complete balanced chemical equation.

#### 2. Solve for the Unknown

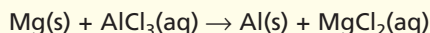
Iron is listed above copper in the metals activity series. Therefore, the first reaction will take place because iron is more reactive than copper. In this case, iron will replace copper. The skeleton equation for this reaction is



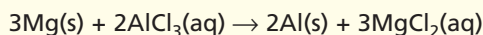
This equation is balanced.

In the second reaction, chlorine is more reactive than bromine because bromine is listed below chlorine in the halogen activity series. Therefore, the reaction will not take place. The skeleton equation for this situation is  $\text{Br(l)} + \text{MgCl}_2\text{(aq)} \rightarrow \text{NR}$ . No balancing is required.

Magnesium is listed above aluminum in the metals activity series. Therefore, the third reaction will take place because magnesium is more reactive than aluminum. In this case, magnesium will replace aluminum. The skeleton equation for this reaction is



This equation is not balanced. The balanced equation is



#### 3. Evaluate the Answer

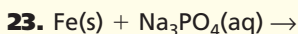
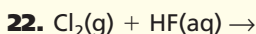
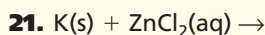
The activity series shown in **Figure 10-10** supports the reaction predictions. The chemical equations are balanced because the number of atoms of each substance is equal on both sides of the equation.



Magnesium is an essential element for the human body. You can ensure an adequate magnesium intake by eating magnesium-rich foods.

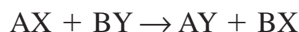
## PRACTICE PROBLEMS

Predict if the following single-replacement reactions will occur. If a reaction occurs, write a balanced equation for the reaction.

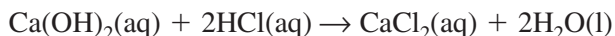


For more practice with predicting if single-replacement reactions will occur, go to **Supplemental Practice Problems** in Appendix A.

**Double-replacement reactions** The final type of replacement reaction which involves an exchange of ions between two compounds is called a **double-replacement reaction**.

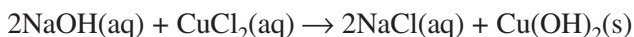


In this generic equation, A and B represent positively charged ions (cations), and X and Y represent negatively charged ions (anions). You can see that the anions have switched places and are now bonded to the other cations in the reaction. In other words, X replaces Y and Y replaces X—a double replacement. More simply, you might say that the positive and negative ions of two compounds switch places. The reaction between calcium hydroxide and hydrochloric acid is a double-replacement reaction.



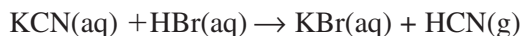
The ionic components of the reaction are  $\text{Ca}^{2+}$ ,  $\text{OH}^-$ ,  $\text{H}^+$ , and  $\text{Cl}^-$ . Knowing this, you can now see the two replacements of the reaction. The anions ( $\text{OH}^-$  and  $\text{Cl}^-$ ) have changed places and are now bonded to the other cations ( $\text{Ca}^{2+}$  and  $\text{H}^+$ ) in the reaction.

The reaction between sodium hydroxide and copper(II) chloride in solution is also a double-replacement reaction.



In this case, the anions ( $\text{OH}^-$  and  $\text{Cl}^-$ ) changed places and are now associated with the other cations ( $\text{Na}^+$  and  $\text{Cu}^{2+}$ ). The result of this reaction is a solid product, copper(II) hydroxide. A solid produced during a chemical reaction in a solution is called a **precipitate**.

One of the key characteristics of double-replacement reactions is the type of product that is formed when the reaction takes place. All double-replacement reactions produce either a precipitate, a gas, or water. An example of a double-replacement reaction that forms a gas is that of potassium cyanide and hydrobromic acid.



It is important to be able to evaluate the chemistry of double-replacement reactions and predict the products of these reactions. The basic steps to do this are given in **Table 10-2**.

**Table 10-2**

Guidelines for Double-Replacement Reactions	
Step	Example
1. Write the components of the reactants in a skeleton equation.	$\text{Al(NO}_3)_3 + \text{H}_2\text{SO}_4$
2. Identify the cations and anions in each compound.	$\text{Al(NO}_3)_3$ has $\text{Al}^{3+}$ and $\text{NO}_3^-$ $\text{H}_2\text{SO}_4$ has $\text{H}^+$ and $\text{SO}_4^{2-}$
3. Pair up each cation with the anion from the other compound.	$\text{Al}^{3+}$ pairs with $\text{SO}_4^{2-}$ $\text{H}^+$ pairs with $\text{NO}_3^-$
4. Write the formulas for the products using the pairs from step 3.	$\text{Al}_2(\text{SO}_4)_3$ $\text{HNO}_3$
5. Write the complete equation for the double-replacement reaction.	$\text{Al(NO}_3)_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{HNO}_3$
6. Balance the equation.	$2\text{Al(NO}_3)_3 + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 6\text{HNO}_3$

Now use this information to work the following practice problems.

## PRACTICE PROBLEMS

Write the balanced chemical equations for the following double-replacement reactions.

24. Aqueous lithium iodide and aqueous silver nitrate react to produce solid silver iodide and aqueous lithium nitrate.
25. Aqueous barium chloride and aqueous potassium carbonate react to produce solid barium carbonate and aqueous potassium chloride.
26. Aqueous sodium oxalate and aqueous lead(II) nitrate react to produce solid lead(II) oxalate and aqueous sodium nitrate.



For more practice with writing double-replacement equations, go to **Supplemental Practice Problems** in Appendix A.

Now that you have learned about the various types of chemical reactions, you can use **Table 10-3** to help you organize them in a way such that you can identify each and predict its products.

As the table indicates, the components of double-replacement reactions are dissolved in water. As you continue with Section 10.3, you will learn more about double-replacement reactions in aqueous solutions.

**Table 10-3**

Predicting Products of Chemical Reactions		
Class of reaction	Reactants	Probable products
Synthesis	Two or more substances	One compound
Combustion	A metal and oxygen A nonmetal and oxygen A compound and oxygen	The oxide of the metal The oxide of the nonmetal Two or more oxides
Decomposition	One compound	Two or more elements and/or compounds
Single-replacement	A metal and a compound A nonmetal and a compound	A new compound and the replaced metal A new compound and the replaced nonmetal
Double-replacement	Two compounds	Two different compounds, one of which is often a solid, water, or a gas

## Section 10.2 Assessment

27. What are the five classes of chemical reactions?
28. Identify two characteristics of combustion reactions.
29. Compare and contrast single-replacement reactions and double-replacement reactions.
30. Describe the result of a double-replacement reaction.
31. **Thinking Critically** Does the following reaction occur? Explain your answer.  
$$3\text{Ni} + 2\text{AuBr}_3 \rightarrow 3\text{NiBr}_2 + 2\text{Au}$$
32. **Classifying** What type of reaction is most likely to occur when barium reacts with fluorine? Write the chemical equation for the reaction.



## Objectives

- **Describe** aqueous solutions.
- **Write** complete ionic and net ionic equations for chemical reactions in aqueous solutions.
- **Predict** whether reactions in aqueous solutions will produce a precipitate, water, or a gas.

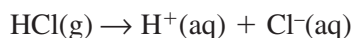
## Vocabulary

solute  
solvent  
aqueous solution  
complete ionic equation  
spectator ion  
net ionic equation

Many of the reactions discussed in the previous section involve substances dissolved in water. When a substance dissolves in water, a solution forms. You learned in Chapter 3 that a solution is a homogeneous mixture. A solution contains one or more substances called **solutes** dissolved in the water. In this case, water is the **solvent**, the most plentiful substance in the solution. An **aqueous solution** is a solution in which the solvent is water. Read the **How It Works** feature at the end of this chapter to see how aqueous solutions are used in hot and cold packs.

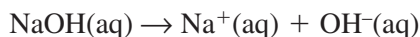
## Aqueous Solutions

Although water is always the solvent in aqueous solutions, there are many possible solutes. Some solutes, such as sucrose (table sugar) and ethanol (grain alcohol), are molecular compounds that exist as molecules in aqueous solutions. Other solutes are molecular compounds that form ions when they dissolve in water. For example, the molecular compound hydrogen chloride forms hydrogen ions and chloride ions when it dissolves in water, as shown in **Figure 10-11**. An equation can be used to show this process.



Compounds such as hydrogen chloride that produce hydrogen ions in aqueous solution are acids. In fact, an aqueous solution of hydrogen chloride is often referred to as hydrochloric acid. You'll learn more about acids in Chapter 19.

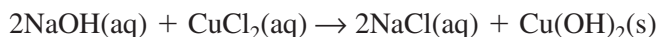
In addition to molecular compounds, ionic compounds may be solutes in aqueous solutions. Recall from Chapter 8 that ionic compounds consist of positive ions and negative ions held together by ionic bonds. When ionic compounds dissolve in water, their ions can separate. The equation below shows an aqueous solution of the ionic compound sodium hydroxide.



When two aqueous solutions that contain ions as solutes are combined, the ions may react with one another. These reactions are always double-replacement reactions. The solvent molecules, which are all water molecules, do not usually react. Three types of products can form from the double-replacement reaction: precipitate, water, or gas. You can observe a precipitate forming when you do the **miniLAB** for this chapter.

## Reactions That Form Precipitates

Some reactions that occur in aqueous solutions produce precipitates. For example, when aqueous solutions of sodium hydroxide and copper(II) chloride are mixed, a double-replacement reaction occurs in which the precipitate copper(II) hydroxide forms.

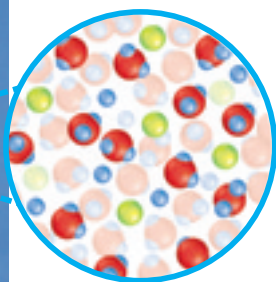


This is shown in **Figure 10-12**.



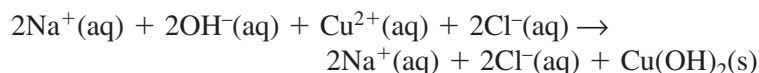
**Figure 10-11**

In an aqueous solution, hydrogen chloride (HCl) breaks apart into hydrogen ions ( $\text{H}^+$ ) and chloride ions ( $\text{Cl}^-$ ).

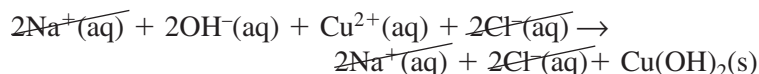


Note that the chemical equation does not show some details of this reaction. Sodium hydroxide and copper(II) chloride are ionic compounds. Therefore, in aqueous solutions they exist as  $\text{Na}^+$ ,  $\text{OH}^-$ ,  $\text{Cu}^{2+}$ , and  $\text{Cl}^-$  ions. When their solutions are combined,  $\text{Cu}^{2+}$  ions in one solution and  $\text{OH}^-$  ions in the other solution react to form the precipitate copper(II) hydroxide,  $\text{Cu}(\text{OH})_2(\text{s})$ . The  $\text{Na}^+$  and  $\text{Cl}^-$  ions remain dissolved in the new solution.

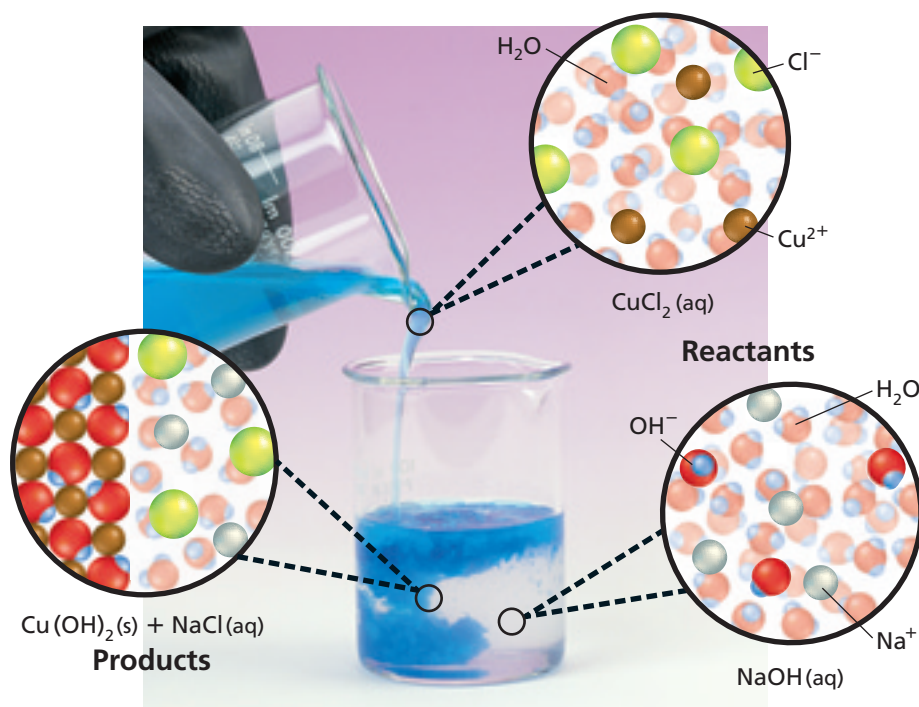
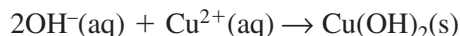
To show the details of reactions that involve ions in aqueous solutions, chemists use ionic equations. Ionic equations differ from chemical equations in that substances that are ions in solution are written as ions in the equation. Look again at the reaction between aqueous solutions of sodium hydroxide and copper(II) chloride. To write the ionic equation for this reaction, you must show the reactants  $\text{NaOH}(\text{aq})$  and  $\text{CuCl}_2(\text{aq})$  and the product  $\text{NaCl}(\text{aq})$  as ions.



An ionic equation that shows all of the particles in a solution as they realistically exist is called a **complete ionic equation**. Note that the sodium ions and the chloride ions are both reactants and products. Because they are both reactants and products, they do not participate in the reaction. Ions that do not participate in a reaction are called **spectator ions** and usually are not shown in ionic equations. Ionic equations that include only the particles that participate in the reaction are called **net ionic equations**. Net ionic equations are written from complete ionic equations by crossing out all spectator ions. For example, a net ionic equation is what remains after the sodium and chloride ions are crossed out of this complete ionic equation.



Only the hydroxide and copper ions are left in the net ionic equation shown below.



**Figure 10-12**

Like the aqueous solution of  $\text{HCl}$  you saw in **Figure 10-11**, sodium hydroxide ( $\text{NaOH}$ ) in an aqueous solution dissociates into its ions sodium ( $\text{Na}^+$ ) and hydroxide ( $\text{OH}^-$ ). Copper chloride ( $\text{CuCl}_2$ ) also dissociates into its ions,  $\text{Cu}^{2+}$  and  $\text{Cl}^-$ .

## EXAMPLE PROBLEM 10-3



Barium carbonate ( $\text{BaCO}_3$ ) is used in the rubber that is eventually processed into tires.

### Reactions That Form a Precipitate

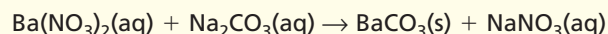
Write the chemical, complete ionic, and net ionic equations for the reaction between aqueous solutions of barium nitrate and sodium carbonate that forms the precipitate barium carbonate.

#### 1. Analyze the Problem

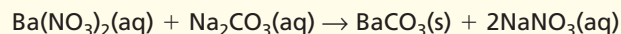
You are given the word equation for the reaction between barium nitrate and sodium carbonate. You must determine the chemical formulas and relative amounts of all reactants and products to write the chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation you can write the net ionic equation. The net ionic equation will include fewer substances than the other equations.

#### 2. Solve for the Unknown

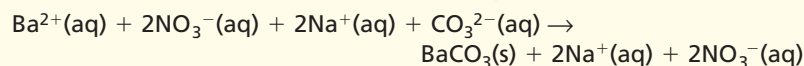
Write the correct chemical formulas and physical states for all substances involved in the reaction.



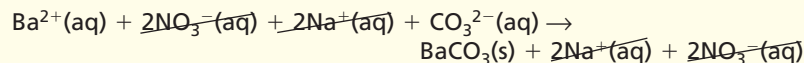
Balance the skeleton equation.



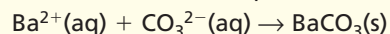
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.



#### 3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only the reacting particles. The particles that compose the solid precipitate that is the result of the reaction are no longer ions.

## PRACTICE PROBLEMS

Write chemical, complete ionic, and net ionic equations for the following reactions that may produce precipitates. Use NR to indicate that no reaction occurs.

33. Aqueous solutions of potassium iodide and silver nitrate are mixed, forming the precipitate silver iodide.
34. Aqueous solutions of ammonium phosphate and sodium sulfate are mixed. No precipitate forms and no gas is produced.
35. Aqueous solutions of aluminum chloride and sodium hydroxide are mixed, forming the precipitate aluminum hydroxide.
36. Aqueous solutions of lithium sulfate and calcium nitrate are mixed, forming the precipitate calcium sulfate.
37. Aqueous solutions of sodium carbonate and manganese(V) chloride are mixed, forming the precipitate manganese(V) carbonate.

Practice!

For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

## Observing a Precipitate-Forming Reaction

**Applying Concepts** When two clear, colorless solutions are mixed, a chemical reaction may occur, resulting in the formation of a precipitate.

**Materials** 150-mL beakers (2); 100-mL graduated cylinder; stirring rod (2); spatula (2); weighing paper (2); NaOH; Epsom salts ( $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ ); distilled water, balance

### Procedure



- CAUTION:** Use gloves when working with NaOH. Measure about 4 g NaOH and place it in a 150-mL beaker. Add 50 mL distilled water to the NaOH. Mix with a stirring rod until the NaOH dissolves.
- Measure about 6 g Epsom salts and place it in another 150-mL beaker. Add 50 mL distilled water to the Epsom salts. Mix with another stirring rod until the Epsom salts dissolve.
- Slowly pour the Epsom salts solution into the NaOH solution. Record your observations.
- Stir the new solution. Record your observations.
- Allow the precipitate to settle, then decant the



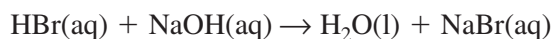
liquid from the solid. Dispose of the solid as your teacher instructs.

### Analysis

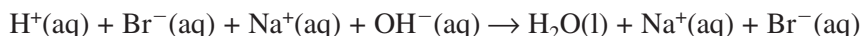
- Write a chemical equation for the reaction between the NaOH and  $\text{MgSO}_4$ . Most sulfate compounds exist as ions in aqueous solutions.
- Write the complete ionic equation for this reaction.
- Write the net ionic equation for this reaction.

## Reactions That Form Water

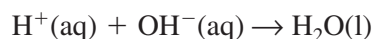
Another type of double-replacement reaction that occurs in an aqueous solution produces water molecules. The water molecules produced in the reaction increase the number of solvent particles. Unlike reactions in which a precipitate forms, no evidence of a chemical reaction is observable because water is colorless and odorless and already makes up most of the solution. For example, when you mix hydrobromic acid with a sodium hydroxide solution, a double-replacement reaction occurs and water is formed.



In this case, the reactants and the product sodium bromide exist as ions in an aqueous solution. The complete ionic equation for this reaction shows these ions.



Look carefully at the complete ionic equation. The reacting solute ions are the hydrogen and hydroxide ions because the sodium and bromine ions are both spectator ions. If you cross out the spectator ions, you are left with the ions that take part in the reaction.



This equation is the net ionic equation for the reaction.



## EXAMPLE PROBLEM 10-4



Hearing aids often use lithium batteries because batteries made with lithium are lightweight and have a long lifetime.

### Reactions That Form Water

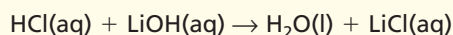
Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous lithium hydroxide, which produces water.

#### 1. Analyze the Problem

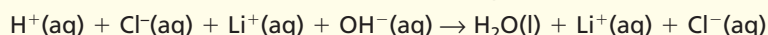
You are given the word equation for the reaction that occurs between hydrochloric acid and lithium hydroxide. You must determine the chemical formulas for and relative amounts of all reactants and products to write the chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation you can write the net ionic equation.

#### 2. Solve for the Unknown

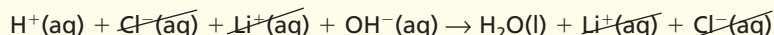
Write the skeleton equation for the reaction and balance it.



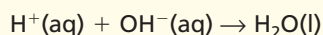
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.



#### 3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces water. The particles that compose the product water are no longer ions.

## PRACTICE PROBLEMS

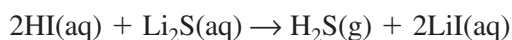
Write chemical, complete ionic, and net ionic equations for the reactions between the following substances, which produce water.

38. Sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and aqueous potassium hydroxide
39. Hydrochloric acid ( $\text{HCl}$ ) and aqueous calcium hydroxide
40. Nitric acid ( $\text{HNO}_3$ ) and aqueous ammonium hydroxide
41. Hydrosulfuric acid ( $\text{H}_2\text{S}$ ) and aqueous calcium hydroxide
42. Phosphoric acid ( $\text{H}_3\text{PO}_4$ ) and aqueous magnesium hydroxide

## Reactions That Form Gases

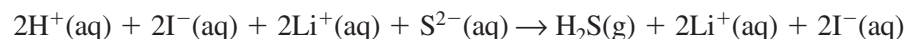
A third type of double-replacement reaction that occurs in aqueous solutions results in the formation of a gas. Some gases commonly produced in these reactions are carbon dioxide, hydrogen cyanide, and hydrogen sulfide.

A gas-producing reaction occurs when you mix hydroiodic acid ( $\text{HI}$ ) with an aqueous solution of lithium sulfide. Bubbles of hydrogen sulfide gas form in the container during the reaction. Lithium iodide is also produced in this reaction and remains dissolved in the solution.

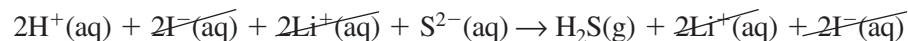


For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

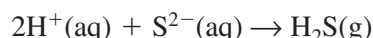
The reactants hydroiodic acid and lithium sulfide exist as ions in aqueous solution. Therefore, you can write an ionic equation for this reaction. The complete ionic equation includes all of the substances in the solution.



Note that there are many spectator ions in the equation. When the spectator ions are crossed out, only the substances involved in the reaction remain in the equation.



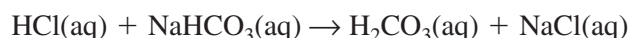
This is the net ionic equation.



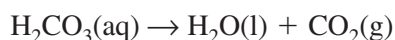
You observed another gas-producing reaction in the **DISCOVERY LAB** at the beginning of this chapter. In that reaction carbon dioxide gas was produced and bubbled out of the solution. Another reaction that produces carbon dioxide gas occurs in your kitchen when you mix vinegar and baking soda. Vinegar is an aqueous solution of acetic acid and water. Baking soda essentially consists of sodium hydrogen carbonate. Rapid bubbling occurs when vinegar and baking soda are combined. The bubbles are carbon dioxide gas escaping from the solution. You can see this reaction occurring in **Figure 10-13**.

A reaction similar to the one between vinegar and baking soda occurs when you combine any acidic solution and sodium hydrogen carbonate. In all cases, two reactions must occur almost simultaneously in the solution to produce the carbon dioxide gas. One reaction is double-replacement and the other is decomposition.

For example, when you dissolve sodium hydrogen carbonate in hydrochloric acid, a gas-producing double-replacement reaction occurs. The hydrogen in the hydrochloric acid and the sodium in the sodium hydrogen carbonate replace each other.



Sodium chloride is an ionic compound and its ions remain separate in the aqueous solution. However, as the carbonic acid ( $\text{H}_2\text{CO}_3$ ) forms, it decomposes immediately into water and carbon dioxide.



## Careers Using Chemistry

### Pastry Chef

*Do you like to make desserts that are as wonderful to look at as they are to eat? Then consider a career as a pastry chef.*

Pastry chefs work in hotels, bakeries, catering, and manufacturing. Like the owner of a business, a pastry chef also must take on financial and managerial tasks. Creativity as well as a solid understanding of the chemistry of cooking is a must.

**Figure 10-13**

When vinegar and baking soda (sodium hydrogen carbonate,  $\text{NaHCO}_3$ ) combine, the result is a vigorous bubbling that releases carbon dioxide ( $\text{CO}_2$ ).

<b>Reaction 1</b>	$\text{HCl(aq)} + \text{NaHCO}_3\text{(aq)} \rightarrow \text{H}_2\text{CO}_3\text{(aq)} + \text{NaCl(aq)}$
<b>Reaction 2</b>	$\text{H}_2\text{CO}_3\text{(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$
<b>Combined equation</b>	$\text{HCl(aq)} + \text{NaHCO}_3\text{(aq)} + \cancel{\text{H}_2\text{CO}_3\text{(aq)}} \rightarrow$ $\cancel{\text{H}_2\text{CO}_3\text{(aq)}} + \text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$
<b>Overall equation</b>	$\text{HCl(aq)} + \text{NaHCO}_3\text{(aq)} \rightarrow$ $\text{H}_2\text{O(l)} + \text{CO}_2\text{(g)} + \text{NaCl(aq)}$

$$\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$$
$$\text{H}^+(\text{aq}) + \cancel{\text{Cl}^-(\text{aq})} + \cancel{\text{Na}^+(\text{aq})} + \text{HCO}_3^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) + \cancel{\text{Na}^+(\text{aq})} + \cancel{\text{Cl}^-(\text{aq})}$$
$$\text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

### Figure 10-14

A collection of household cleaning products is displayed against a blue background. The items include a large yellow box of baking soda with a red circular logo that says "ODOR ABSORBING BAKING SODA", a red spray bottle, a white container of odor absorber, and a box of disinfectant. There is also a small box of "NEW" disinfectant and a box of "Baking Soda &amp; Peroxide".

## EXAMPLE PROBLEM 10-5

### Reactions That Form Gases

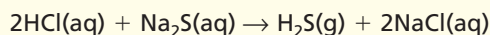
Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous sodium sulfide, which produces hydrogen sulfide gas.

#### 1. Analyze the Problem

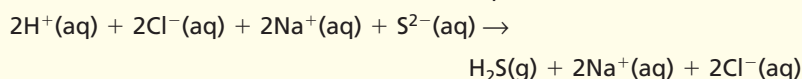
You are given the word equation for the reaction between hydrochloric acid and sodium sulfide. You must write the skeleton equation and balance it. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions in the complete ionic equation, you can write the net ionic equation.

#### 2. Solve for the Unknown

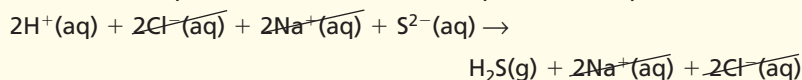
Write the correct skeleton equation for the reaction and balance it.



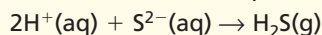
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation in its smallest whole number ratio.



#### 3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces hydrogen sulfide. The particles that compose the product are no longer ions.



Many chemical reactions are obvious by the results that can be detected by the senses. The chemical reaction that occurs when eggs rot gives off hydrogen sulfide ( $\text{H}_2\text{S}$ ) gas, which has a pungent, distinctive odor.

## PRACTICE PROBLEMS

Write chemical, complete ionic, and net ionic equations for these reactions.

43. Perchloric acid ( $\text{HClO}_4$ ) reacts with aqueous potassium carbonate.
44. Sulfuric acid ( $\text{H}_2\text{SO}_4$ ) reacts with aqueous sodium cyanide.
45. Hydrobromic acid ( $\text{HBr}$ ) reacts with aqueous ammonium carbonate.
46. Nitric acid ( $\text{HNO}_3$ ) reacts with aqueous potassium rubidium sulfide.



For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

## Section 10.3 Assessment

47. Describe an aqueous solution.
48. Distinguish between a complete ionic equation and a net ionic equation.
49. What are three common types of products produced by reactions that occur in aqueous solutions?
50. **Thinking Critically** Explain why net ionic equations communicate more than chemical equations about reactions in aqueous solutions.
51. **Communicating** Describe the reaction of aqueous solutions of sodium sulfide and copper(II) sulfate, producing the precipitate copper(II) sulfide.



## Activities of Metals

Some metals are more reactive than others. By comparing how different metals react with the same ions in aqueous solutions, an activity series for the tested metals can be developed. The activity series will reflect the relative reactivity of the tested metals. It can be used to predict whether reactions will occur.

### Problem

Which is the most reactive metal tested? Which is the least reactive metal tested? Can this information be used to predict whether reactions will occur?

### Objectives

- **Observe** chemical reactions.
- **Sequence** the activities of some metals.
- **Predict** if reactions will occur between certain substances.

### Materials

1.0M  $\text{Zn}(\text{NO}_3)_2$   
 1.0M  $\text{Al}(\text{NO}_3)_3$   
 1.0M  $\text{Cu}(\text{NO}_3)_2$   
 1.0M  $\text{Mg}(\text{NO}_3)_2$   
 pipettes (4)  
 wire cutters  
 Cu wire

Al wire  
 Mg ribbon  
 Zn metal strips (4)  
 emery cloth or fine sandpaper  
 24-well microscale reaction plate

### Safety Precautions



- Always wear safety goggles and a lab apron.
- Use caution when using sharp and coarse equipment.

### Pre-Lab

1. Read the entire CHEMLAB.
2. Make notes about procedures and safety precautions to use in the laboratory.
3. Prepare your data table.

Reactions Between Solutions and Metals

	$\text{Al}(\text{NO}_3)_3$	$\text{Mg}(\text{NO}_3)_2$	$\text{Zn}(\text{NO}_3)_2$	$\text{Cu}(\text{NO}_3)_2$
Al				
Mg				
Zn				
Cu				

4. Form a hypothesis about what reactions will occur.
5. What are the independent and dependent variables?
6. What gas is produced when magnesium and hydrochloric acid react? Write the chemical equation for the reaction.
7. Why is it important to clean the magnesium ribbon? How might not polishing a piece of metal affect the reaction involving that metal?

### Procedure

1. Use a pipette to fill each of the four wells in column 1 of the reaction plate with 2 mL of 1.0M  $\text{Al}(\text{NO}_3)_3$  solution.
2. Repeat the procedure in step 1 to fill the four wells in column 2 with 2 mL of 1.0M  $\text{Mg}(\text{NO}_3)_2$  solution.
3. Repeat the procedure in step 1 to fill the four wells in column 3 with 2 mL of 1.0M  $\text{Zn}(\text{NO}_3)_2$  solution.
4. Repeat the procedure in step 1 to fill the four wells in column 4 with 2 mL of 1.0M  $\text{Cu}(\text{NO}_3)_2$  solution.
5. With the emery paper or sandpaper, polish 10 cm of aluminum wire until it is shiny. Use wire cutters to cut the lead wire into four 2.5-cm pieces. Place a piece of the aluminum wire in each row A well that contains solution.
6. Repeat the procedure in step 5 using 10 cm of magnesium ribbon. Place a piece of the Mg ribbon in each row B well that contains solution.
7. Use the emery paper or sandpaper to polish small strips of zinc metal. Place a piece of Zn metal in each row C well that contains solution.



8. Repeat the procedure in step 5 using 10 cm of copper wire. Place a piece of Cu wire in each row D well that contains solution.
9. Observe what happens in each cell. After five minutes, record your observations on the data table you made.

### Cleanup and Disposal

1. Dispose of all chemicals and solutions as directed by your teacher.
2. Clean your equipment and return it to its proper place.
3. Wash your hands thoroughly before you leave the lab.

### Analyze and Conclude

1. **Observing and Inferring** In which wells of the reaction plate did chemical reactions occur? Which metal reacted with the most solutions? Which metal reacted with the fewest solutions? Which metal is the most reactive?
2. **Sequencing** The most active metal reacted with the most solutions. The least active metal reacted with the fewest solutions. Order the four metals from the most active to the least active.
3. **Comparing and Contrasting** Compare your activity series with the activity series shown here. How does the order you determined for the four metals you tested compare with the order of these metals?

4. **Applying Concepts** Write a chemical equation for each single-replacement reaction that occurred on your reaction plate.
5. **Predicting** Use the diagram below to predict if a single-replacement reaction will occur between the following reactants. Write a chemical equation for each reaction that will occur.
  - a. Ca and  $\text{Sn}(\text{NO}_3)_2$
  - b. Ag and  $\text{Ni}(\text{NO}_3)_2$
  - c. Cu and  $\text{Pb}(\text{NO}_3)_3$
6. **Error Analysis** If the activity series you sequenced does not agree with the order in the diagram below, propose a reason for the disagreement.

Most active	↓	<b>METALS</b>
		Lithium
		Rubidium
		Potassium
		Calcium
		Sodium
		Magnesium
		Aluminum
		Manganese
		Zinc
		Iron
		Nickel
		Tin
		Lead
		Copper
		Silver
		Platinum
		Gold
Least active	↓	
Most active	↓	<b>HALOGENS</b>
		Fluorine
		Chlorine
		Bromine
Least active	↓	Iodine

### Real-World Chemistry

1. Under what circumstances might it be important to know the activity tendencies of a series of elements?
2. Describe some of the environmental impacts of nitrates.

# How It Works

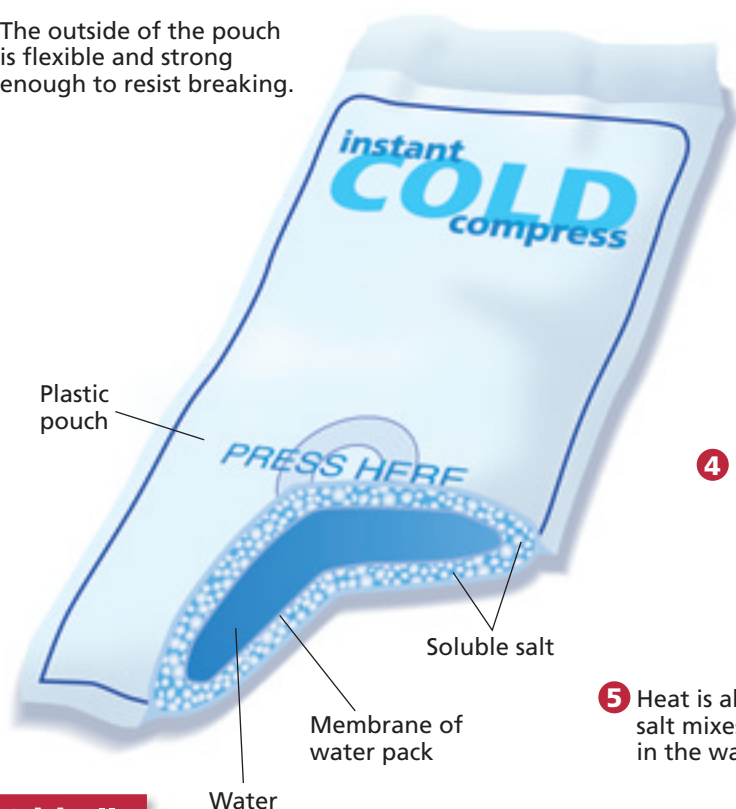
## Hot and Cold Packs

Athletes know that the application of heat or cold to a strain or sprain usually relieves the pain and may lessen the severity of the injury. Instant hot and cold packs allow you to quickly and easily apply the appropriate remedy to the injury.

Hot and cold packs create aqueous solutions of a soluble salt. A salt such as ammonium nitrate is used in the cold pack and heat is absorbed as the salt dissolves in the water. Hot packs release heat when a salt such as calcium chloride dissolves in the water.



- 1 The outside of the pouch is flexible and strong enough to resist breaking.



- 2 The water is stored in an inner chamber made from thin, easily broken plastic membrane.

- 3 Salt is stored between the outer covering and the water container.

- 4 The water is released from the inner chamber when you squeeze or strike the pack.

- 5 Heat is absorbed as the salt mixes with and dissolves in the water.

### Thinking Critically

- Predicting** An aqueous solution of sodium thiosulfate releases heat when it crystallizes. What would happen when sodium thiosulfate crystals dissolve in water?
- Hypothesizing** One type of heat pack contains fine iron particles. These packs are kept in a sealed container and release heat when they are exposed to air. How does this type of pack work?

## Summary

### 10.1 Reactions and Equations

- Some chemical reactions release energy in the form of heat and light, and some absorb energy.
- Changes in temperature, color, odor, and physical state are all types of evidence that indicate a chemical reaction has occurred.
- Word and skeleton equations provide important information about a chemical reaction, such as the reactants and products involved in the reaction and their physical states.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction. Chemical equations are balanced.
- Balancing an equation involves adjusting the coefficients of the chemical formulas in the skeleton equation until the number of atoms of each element is equal on both sides of the equation.

### 10.2 Classifying Chemical Reactions

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Synthesis, combustion, decomposition, single-replacement, and double-replacement reactions are five classes of chemical reactions.
- A synthesis reaction occurs when two substances react to yield a single product. The substances that react can be two elements, a compound and an element, or two compounds.
- A combustion reaction occurs when a substance reacts with oxygen, producing heat and light.
- A decomposition reaction occurs when a single compound breaks down into two or more elements or new compounds.

- A single-replacement reaction occurs when the atoms of one element replace the atoms of another element in a compound.
- In single-replacement reactions, a metal may replace hydrogen in water, a metal may replace another metal in a compound dissolved in water, and a non-metal may replace another nonmetal in a compound.
- Metals and halogens can be ordered according to their reactivities. These listings, which are called activity series, can be used to predict if single-replacement reactions will occur.
- A double-replacement reaction involves the exchange of positive ions between two compounds.

### 10.3 Reactions in Aqueous Solutions

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When most ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions may react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.
- Three types of products produced during reactions in aqueous solutions are precipitates, water, and gases.
- An ionic equation shows the details of reactions in aqueous solutions. A complete ionic equation shows all the particles in a solution as they exist. A net ionic equation includes only the particles that participate in a reaction in a solution.

## Vocabulary

- |                                    |  |  |
|------------------------------------|--|--|
| • aqueous solution (p. 292)        | • decomposition reaction (p. 286)      | • reactant (p. 278)                    |
| • chemical equation (p. 280)       | • double-replacement reaction (p. 290) | • single-replacement reaction (p. 287) |
| • chemical reaction (p. 277)       | • net ionic equation (p. 293)          | • solute (p. 292)                      |
| • coefficient (p. 280)             | • precipitate (p. 290)                 | • solvent (p. 292)                     |
| • combustion reaction (p. 285)     | • product (p. 278)                     | • spectator ion (p. 293)               |
| • complete ionic equation (p. 293) |  | • synthesis reaction (p. 284)          |



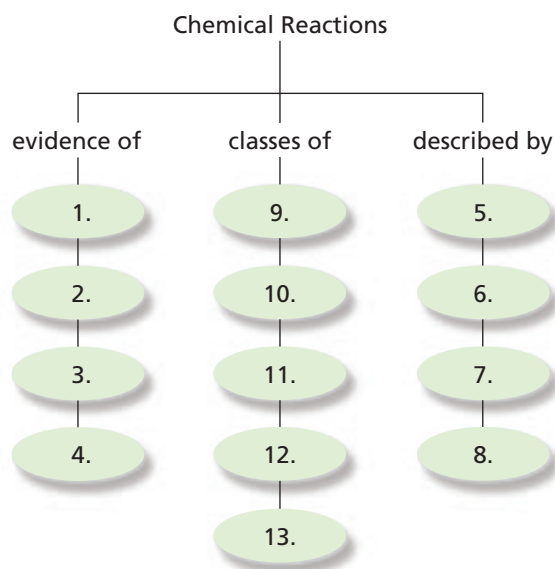


CLICK HERE

Go to the Chemistry Web site at [science.glencoe.com](http://science.glencoe.com) or use the Chemistry CD-ROM for additional Chapter 10 Assessment.

## Concept Mapping

52. Use the following terms and phrases to complete the concept map: synthesis, net ionic equation, change in energy, change in physical state, single-replacement, word equation, decomposition, complete ionic equation, double-replacement, combustion, change in odor, chemical equation, change in color.



## Mastering Concepts

53. Explain the difference between reactants and products. (10.1)
54. What do the arrows and coefficients used by chemists in equations communicate? (10.1)
55. Write formulas for the following substances and designate their physical states. (10.1)
- nitrogen dioxide gas
  - liquid gallium
  - barium chloride dissolved in water
  - solid ammonium carbonate
56. Identify the reactants in the following reaction: When potassium is dropped into aqueous zinc nitrate, zinc and aqueous potassium nitrate form. (10.1)

57. When gasoline is burned in an automobile engine, what evidence indicates that a chemical change has occurred? (10.1)
58. Write the word equation for this skeleton equation. (10.1)
- $$\text{Mg(s)} + \text{FeCl}_3\text{(aq)} \rightarrow \text{Fe(s)} + \text{MgCl}_2\text{(aq)}$$
59. Balance the equation in question 58. (10.1)
60. What are five classes of chemical reactions? (10.2)
61. How would you classify a chemical reaction between two reactants that produces one product? (10.2)
62. Explain the difference between a single-replacement reaction and a double-replacement reaction. (10.2)
63. Under what conditions does a precipitate form in a chemical reaction? (10.2)
64. Classify the chemical reaction in question 58. (10.2)
65. In each of the following pairs, which element will replace the other in a reaction? (10.2)
- tin and sodium
  - fluorine and iodine
  - lead and silver
  - copper and nickel
66. When reactions occur in aqueous solutions what common types of products are produced? (10.3)
67. Compare and contrast chemical equations and ionic equations. (10.3)
68. What is a net ionic equation? How does it differ from a complete ionic equation? (10.3)
69. Define spectator ion. (10.3)
70. Write the net ionic equation for a chemical reaction that occurs in an aqueous solution and produces water. (10.3)

## Mastering Problems

### Balancing Chemical Equations (10.1)

71. Write skeleton equations for these reactions.
- hydrogen iodide(g)  $\rightarrow$  hydrogen(g) + iodine(g)
  - aluminum(s) + iodine(s)  $\rightarrow$  aluminum iodide(s)
  - iron(II) oxide(s) + oxygen(g)  $\rightarrow$  iron(III) oxide(s)
72. Write skeleton equations for these reactions.
- butane (C<sub>4</sub>H<sub>10</sub>)(l) + oxygen(g)  $\rightarrow$  carbon dioxide(g) + water(l)
  - aluminum carbonate(s)  $\rightarrow$  aluminum oxide(s) + carbon dioxide(g)
  - silver nitrate(aq) + sodium sulfide(aq)  $\rightarrow$  silver sulfide(s) + sodium nitrate(aq)

- 73.** Write skeleton equations for these reactions.
- iron(s) + fluorine(g) → iron(III) fluoride(s)
  - sulfur trioxide(g) + water(l) → sulfuric acid(aq)
  - sodium(s) + magnesium iodide(aq) → sodium iodide(aq) + magnesium(s)
  - vanadium(s) + oxygen(g) → vanadium(V) oxide(s)
- 74.** Write skeleton equations for these reactions.
- lithium(s) + gold(III) chloride(aq) → lithium chloride(aq) + gold(s)
  - iron(s) + tin(IV) nitrate(aq) → iron(III) nitrate(aq) + tin(s)
  - nickel(II) chloride(s) + oxygen(g) → nickel(II) oxide(s) + dichlorine pentoxide(g)
  - lithium chromate(aq) + barium chloride(aq) → lithium chloride(aq) + barium chromate(s)
- 75.** Balance the skeleton equations for the reactions described in question 71.
- 76.** Balance the skeleton equations for the reactions described in question 72.
- 77.** Balance the skeleton equations for the reactions described in question 73.
- 78.** Balance the skeleton equations for the reactions described in question 74.
- 79.** Write chemical equations for these reactions.
- When solid naphthalene ( $C_{10}H_8$ ) burns in air, the products are gaseous carbon dioxide and liquid water.
  - Bubbling hydrogen sulfide gas through manganese(II) chloride dissolved in water results in the formation of the precipitate manganese(II) sulfide and hydrochloric acid.
  - Solid magnesium reacts with nitrogen gas to produce solid magnesium nitride.
  - Heating oxygen difluoride gas yields oxygen gas and fluorine gas.
- boron + fluorine →
  - germanium + sulfur →
  - zirconium + nitrogen →
  - tetraphosphorus decoxide + water → phosphoric acid
- 86.** Write a chemical equation for the combustion of each of the following substances. If a compound contains the elements carbon and hydrogen, assume that carbon dioxide gas and liquid water are produced.
- solid barium
  - solid boron
  - liquid acetone ( $C_3H_6O$ )
  - liquid octane ( $C_8H_{18}$ )
- 87.** Write chemical equations for each of the following decomposition reactions. One or more products may be identified.
- magnesium bromide →
  - cobalt(II) oxide →
  - titanium(IV) hydroxide → titanium(IV) oxide + water
  - barium carbonate → barium oxide + carbon dioxide
- 88.** Write chemical equations for the following single-replacement reactions that may occur in water. If no reaction occurs, write NR in place of the products.
- nickel + magnesium chloride →
  - calcium + copper(II) bromide →
  - potassium + aluminum nitrate →
  - magnesium + silver nitrate →
- 89.** Write chemical equations for each of the following double-replacement reactions that occur in water.
- rubidium iodide + silver nitrate →
  - sodium phosphate + manganese(II) chloride →
  - lithium carbonate + molybdenum(VI) bromide →
  - calcium nitrate + aluminum hydroxide →

## Classifying Chemical Reactions (10.2)

- Classify each of the reactions represented by the chemical equations in question 75.
- Classify each of the reactions represented by the chemical equations in question 76.
- Classify each of the reactions represented by the chemical equations in question 77.
- Classify each of the reactions represented by the chemical equations in question 78.
- Classify each of the reactions represented by the chemical equations in question 79.
- Write chemical equations for each of the following synthesis reactions.

## Reactions in Aqueous Solutions (10.3)

- Write complete ionic and net ionic equations for each of the following reactions.
  - $K_2S(aq) + CoCl_2(aq) \rightarrow 2KCl(aq) + CoS(s)$
  - $H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaSO_4(s)$
  - $2HClO(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(ClO)_2(aq)$
- A reaction occurs when hydrosulfuric acid ( $H_2S$ ) is mixed with an aqueous solution of iron(III) bromide. Solid iron(III) sulfide is produced. Write the chemical and net ionic equations for the reaction.

**92.** Write complete ionic and net ionic equations for each of the following reactions.

- $\text{H}_3\text{PO}_4(\text{aq}) + 3\text{RbOH}(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \text{Rb}_3\text{PO}_4(\text{aq})$
- $\text{HCl}(\text{aq}) + \text{NH}_4\text{OH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NH}_4\text{Cl}(\text{aq})$
- $2\text{HI} + (\text{NH}_4)_2\text{S}(\text{aq}) \rightarrow \text{H}_2\text{S}(\text{g}) + 2\text{NH}_4\text{I}(\text{aq})$
- $\text{HNO}_3(\text{aq}) + \text{KCN}(\text{aq}) \rightarrow \text{HCN}(\text{g}) + \text{KNO}_3(\text{aq})$

**93.** A reaction occurs when sulfurous acid ( $\text{H}_2\text{SO}_3$ ) is mixed with an aqueous solution of sodium hydroxide. Aqueous sodium sulfite is produced. Write the chemical and net ionic equations for the reaction.

**94.** A reaction occurs when nitric acid ( $\text{HNO}_3$ ) is mixed with an aqueous solution of potassium hydrogen carbonate. Aqueous potassium nitrate is produced. Write the chemical and net ionic equations for the reaction.

## Mixed Review

*Sharpen your problem-solving skills by answering the following.*

- Identify the products in the following reaction that occurs in plants: Carbon dioxide and water react to produce glucose and oxygen.
- How will aqueous solutions of sucrose and hydrogen chloride differ?
- Write the word equation for each of these skeleton equations.  $\text{C}_6\text{H}_6$  is the formula for benzene.
  - $\text{C}_6\text{H}_6(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
  - $\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
- Write skeleton equations for the following reactions.
  - ammonium phosphate(aq) + chromium(III) bromide(aq)  $\rightarrow$  ammonium bromide(aq) + chromium(III) phosphate(s)
  - chromium(VI) hydroxide(s)  $\rightarrow$  chromium(VI) oxide(s) + water(l)
  - aluminum(s) + copper(I) chloride(aq)  $\rightarrow$  aluminum chloride(aq) + copper(s)
  - potassium iodide(aq) + mercury(I) nitrate(aq)  $\rightarrow$  potassium nitrate(aq) + mercury(I) iodide(s)
- Balance the skeleton equations for the reactions described in question 98.
- Classify each of the reactions represented by the chemical equations in question 99.
- Predict whether each of the following reactions will occur in aqueous solutions. If you predict that a reaction will not occur, explain your reasoning. Note: Barium sulfate and silver bromide precipitate in aqueous solutions.

- sodium hydroxide + ammonium sulfate  $\rightarrow$
- niobium(V) sulfate + barium nitrate  $\rightarrow$
- strontium bromide + silver nitrate  $\rightarrow$

## Thinking Critically

**102. Predicting** A piece of aluminum metal is placed in aqueous KCl. Another piece of aluminum is placed in an aqueous  $\text{AgNO}_3$  solution. Explain why a chemical reaction does or does not occur in each instance.

**103. Designing an Experiment** You suspect that the water in a lake close to your school may contain lead in the form of  $\text{Pb}^{2+}(\text{aq})$  ions. Formulate your suspicion as a hypothesis and design an experiment to test your theory. Write the net ionic equations for the reactions of your experiment. (Hint: In aqueous solution,  $\text{Pb}^{2+}$  forms compounds that are solids with  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ , and  $\text{SO}_4^{2-}$  ions.)

**104. Applying Concepts** Write the chemical equations and net ionic equations for each of the following reactions that may occur in aqueous solutions. If a reaction does not occur, write NR in place of the products. Magnesium phosphate precipitates in an aqueous solution.

- $\text{KNO}_3 + \text{CsCl} \rightarrow$
- $\text{Ca}(\text{OH})_2 + \text{KCN} \rightarrow$
- $\text{Li}_3\text{PO}_4 + \text{MgSO}_4 \rightarrow$
- $\text{HBrO} + \text{NaOH} \rightarrow$

## Writing in Chemistry

- Prepare a poster describing types of chemical reactions that occur in the kitchen.
- Write a report in which you compare and contrast chemical and mathematical equations.

## Cumulative Review

*Refresh your understanding of previous chapters by answering the following.*

- Distinguish among a mixture, a solution, and a compound. (Chapter 3)
- Write the formula for the compounds made from each of the following pairs of ions. (Chapter 9)
  - copper(I) and sulfite
  - tin(IV) and fluoride
  - gold(III) and cyanide
  - lead(II) and sulfide

Use these questions and the test-taking tip to prepare for your standardized test.

- Potassium chromate and lead(II) acetate are both dissolved in a beaker of water, where they react to form solid lead(II) chromate. What is the balanced net ionic equation describing this reaction?
  - $\text{Pb}^{2+}(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2(\text{s})$
  - $\text{Pb}^{2+}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{Pb}(\text{CrO}_4)_2(\text{s})$
  - $\text{Pb}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{PbCrO}_4(\text{s})$
  - $\text{Pb}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{PbC}_2\text{H}_3\text{O}_2(\text{s})$
- What type of reaction is described by the following equation?
 
$$\text{Cs}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CsOH}(\text{aq}) + \text{H}_2(\text{g})$$
  - synthesis
  - combustion
  - decomposition
  - replacement
- Which of the following reactions between halogens and halide salts will occur?
  - $\text{F}_2(\text{g}) + \text{FeI}_2(\text{aq}) \rightarrow \text{FeF}_2(\text{aq}) + \text{I}_2(\text{l})$
  - $\text{I}_2(\text{s}) + \text{MnBr}_2(\text{aq}) \rightarrow \text{MnI}_2(\text{aq}) + \text{Br}_2(\text{g})$
  - $\text{Cl}_2(\text{s}) + \text{SrF}_2(\text{aq}) \rightarrow \text{SrCl}_2(\text{aq}) + \text{F}_2(\text{g})$
  - $\text{Br}_2(\text{l}) + \text{CoCl}_2(\text{aq}) \rightarrow \text{CoBr}_2(\text{aq}) + \text{Cl}_2(\text{g})$

**Interpreting Tables** Use the table to answer questions 4–6.

**Physical Properties of Select Ionic Compounds**

Compound	Name	Physical state at room temp.	Soluble in water?	Melting point (°C)
$\text{NaClO}_3$	sodium chlorate	solid	yes	248
$\text{Na}_2\text{SO}_4$	sodium sulfate	solid	yes	884
$\text{NiCl}_2$	nickel(II) chloride	solid	yes	1009
$\text{Ni}(\text{OH})_2$	nickel(II) hydroxide	solid	no	230
$\text{AgNO}_3$	silver nitrate	solid	yes	212

- An aqueous solution of nickel(II) sulfate is mixed with aqueous sodium hydroxide. Will a visible reaction occur?
  - No, solid nickel(II) hydroxide is soluble in water.
  - No, solid sodium sulfate is soluble in water.

- Yes, solid sodium sulfate will precipitate out of solution.
  - Yes, solid nickel(II) hydroxide will precipitate out of solution.
- When  $\text{AgClO}_3(\text{aq})$  and  $\text{NaNO}_3(\text{aq})$  are mixed, \_\_\_\_\_.
    - no visible reaction occurs
    - solid  $\text{NaClO}_3$  precipitates out of solution
    - $\text{NO}_2$  gas is released from the reaction
    - solid Ag metal is produced
  - Finely ground nickel(II) hydroxide is placed in a beaker of water. It sinks to the bottom of the beaker and remains unchanged. An aqueous solution of hydrochloric acid (HCl) is then added the beaker, and the  $\text{Ni}(\text{OH})_2$  disappears. Which of the following equations best describes what occurred in the beaker?
    - $\text{Ni}(\text{OH})_2(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{NiO}(\text{aq}) + \text{H}_2(\text{g}) + \text{HCl}(\text{aq})$
    - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{NiCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
    - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NiCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
    - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NiCl}_2(\text{aq}) + 3\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
  - The combustion of ethanol,  $\text{C}_2\text{H}_6\text{O}$ , produces carbon dioxide and water vapor. The equation that best describes this process is \_\_\_\_\_.
    - $\text{C}_2\text{H}_6\text{O}(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
    - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
    - $\text{C}_2\text{H}_6\text{O}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$
    - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 3\text{O}_2(\text{l}) + 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
  - What is the product of this synthesis reaction?
 
$$\text{Cl}_2(\text{g}) + 2\text{NO}(\text{g}) \rightarrow ?$$
    - $\text{NCl}_2$
    - $2\text{NOCl}$
    - $\text{N}_2\text{O}_2$
    - $2\text{ClO}$



### TEST-TAKING TIP

**Tables** If a test question involves a table, skim the table before reading the question. Read the title, column heads, and row heads. Then read the question and interpret the information in the table.